Practical Course

Of

Physical chemistry

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Contents:

Experiment (1): Variation of conductance with concentration (strong &weak) electrolytes.

Experiment (2) : Determine Hydrolysis constant by a conductimetric method.

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Experiment (4): Dissociation constant of the weak acid (Conductometric method).

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Some basic concepts

In experiment of chemistry, mostly students require standard solution to study physical properties.

Standard solution:

A solution of known concentration is called standard solution. It contains a known weight of substance in a definite volume of solution. Generally concentrations are expressed in terms of

i) **Normality:** It is the number of gram equivalents of solute present in one liter of solution.

$$\text{Normality (N)} = \frac{\text{Number of gram equivalents}}{\text{Liter of solution}}$$

ii) **Molarity:** It is the number of moles of solute present in one liter of solution.

$$\text{Molarity (M)} = \frac{\text{Number of moles}}{\text{Liter of solution}}$$

iii) **Molality:** It is the number of moles of solute dissolved in 1000 gms of solvent.

$$\text{Molality (m)} = \frac{\text{Number of moles of solute}}{\text{Wt. of solvent in kg (1000 gms)}}$$

iv) **Mole fraction:** The mole fraction of a component in a solution is defined as the number of moles of that component divided by the total number of moles of all components in the solution. If a solute is dissolved in solvent then

$$\text{The mole fraction of solute} = \frac{\text{Moles of solute}}{\text{Moles of solute + Moles of solvent}}$$

$$\text{The mole fraction of solvent} = \frac{\text{Moles of solvent}}{\text{Moles of solute + Moles of solvent}}$$

v) **Volume fraction:** If two liquids are miscible to form a solution then the volume fraction can be determined as follows:

$$\text{Volume fraction of one } (V_1) = \frac{v_1}{v_1 + v_2}$$

$$\text{Volume fraction of other } (V_2) = \frac{v_2}{v_1 + v_2}$$

vi) **Weight fraction:** It is the ratio of weight of solute to the total weight of solution.

$$\text{Wt. fraction of solute} = \frac{\text{Wt. of solute}}{\text{Total wt. of solution}}$$

$$\text{Wt. fraction of solvent} = \frac{\text{Wt. of solvent}}{\text{Total wt. of solution}}$$

If weight fraction multiplied by 100 it is termed as percentage by weight.

vii) **Percentage weight of solute by volume of solvent:** It is the weight of solute dissolved in 100 ml of solution: (Wt/V) and (Wt/Wt). also there is percentage volume of solute by volume of solvent (V/V).

viii) **Parts per thousand (P.P.T.):** This is a method of expression concentration of solute if solution is sufficiently dilute.

If 20 gm of can sugar is dissolved in 1000 gm of water (or 1000 ml) then the resultant solution is said to be 20P.P.T. it can be expressed as

P.P.T = gm/1000 gm of solution.

ix) **Parts per thousand (P.P.m.):** This is the best method of express concentration of solute when solution is expected to be very dilute.

If 20 gm of can sugar is dissolved in one million gm of water (or 1000 ml) then the resultant solution is called 20 P.P.m. it can be expressed as

P.P.m = gm/1000000 gm of solution.

x) **Strength of solution:** The strength of solution is always expressed in terms of grams per liter.

If the concentration is given in terms of normality then

Strength = Normality x Equivalent wt.

And if the concentration is given in terms of normality then

Strength = Molarity x Molecular wt.

xi) **Dilution:** A solution is diluted when extra solvent is added. The concentration of dilute solution can be calculated by using this equation which known as Dilution law

$M_1 V_1 = M_2 V_2$ Or $N_1V_1 = N_2V_2$.

Preparation of solution

1) Preparation solution from solid material

$$Wt_{(gm)} = \frac{M \times V_{ml} \times M.wt}{1000}$$

For example:

prepare o.1 M of NaOH in 100 ml of distilled water.

M.wt of NaOH = 23+16+1= 40 gm/mol

V = 100 ml

M = 0.1 M

Wt = ?

$$Wt_{(gm)} = \frac{M \times V_{ml} \times M.wt}{1000} = \frac{0.1 \times 100 \times 40}{1000} = 0.4 \text{ gm}$$

Take 0.4 gm of NaOH and dissolved in quantity of water, when it dissolved completely add it into 100 ml standard flask and fill it up to the mark.

For example:

Prepare o.1 N of $Ca(OH)_2$ in 100 ml of distilled water.

Eq.wt of $Ca(OH)_2$ = M.wt / 2 = 74/2 = 37 gm / mol , Eq.wt = M.wt / $n_{(no.of\ H+,OH-\ and\ e-)}$

V = 100 ml

N = 0.1 N

Wt = ?

$$Wt_{(gm)} = \frac{N \times V_{ml} \times Eq.wt}{1000} = \frac{0.1 \times 100 \times 37}{1000} = 0.37\ gm$$

Take 0.37 gm of $Ca(OH)_2$ and dissolved in quantity of water, when it dissolved completely add it into 100 ml standard flask and fill it up to the mark.

2) Preparation solution from liquid material

For preparation of dilute solution from concentrate liquid can be use these information. Density, % and M.wt or Eq.wt.

$$Sp.gr = \frac{\text{Density of liquid}}{\text{Density of water at 4°C}}$$

Density of water at 4°C = 1, i.e. Sp.gr = Density

% is the percentage of concentrate liquid in bottle.

$$M = \frac{\text{Sp.gr} \times \% \times 1000}{\text{M.wt}} \quad \text{Or} \quad N = \frac{\text{Sp.gr} \times \% \times 1000}{\text{Eq.wt}}$$

By using dilution law:

$M_1 V_1 = M_2 V_2$ Or $N_1 V_1 = N_2 V_2$

For example:

Prepare 0.5 M and N of H_2SO_4 in 100 ml of distilled water. If you know density of concentrate H_2SO_4 = 1.84 gm/ ml, M.wt= 98 gm/mol and % = 96.

$$M = \frac{\text{Sp.gr} \times \% \times 1000}{\text{M.wt}} \text{ for } N = \frac{\text{Sp.gr} \times \% \times 1000}{\text{Eq.wt}}$$

$$= \frac{1.84 \times 96/100 \times 1000}{98} \text{ for } = \frac{1.84 \times 96/100 \times 1000}{98/2}$$

$$= \frac{1.84 \times 0.96 \times 1000}{98} \quad \text{for} \quad = \frac{1.84 \times 0.96 \times 1000}{49}$$

M = 18.02 for N = 36.04

Now, for Prepare 0.5 M and N of H_2SO_4 in 100 ml of distilled water, using dilution law.

$M_1 V_{1(conc)} = M_2 V_{2(dilute)}$ for $N_1 V_{1(conc)} = N_2 V_{2\,(dilute)}$

$18.02 \times V_{1(conc)} = 0.5 \times 100$ for $36.04 \times V_{1(conc)} = 0.5 \times 100$

$V_{1(conc)} = 2.77$ ml $V_{1(conc)} = 1.38$ ml

Take calculated volume of concentrate liquid and put it in standard flask fill up to the mark by distilled water.

Electrochemistry

Strong electrolytes:

According to Arrhenius theory, strong electrolytes are completely dissociated at any concentration. It means if a strong electrolyte is dissolved in water, it will ionize completely into cations and anions. In other words it can be started that the ionization is 100% and question of calculating degree of dissociation and dissociation constant, does not rise at all. Therefore, Ostwald's law of dilution cannot be made applicable to the strong electrolytes.

If concentrated solution of strong electrolyte is diluted the ions get more space to movement and therefore conductivity of strong electrolyte increases with dilution up to certain extent and attains limiting value.

Electrolytes like HCl, NaOH, KCl have high value of conductance even at low dilution.

Weak electrolytes:

Weak electrolytes are partially dissociated in concentrated solution. The extent of dissociation is called degree of dissociation. The degree of dissociation increases with dilution and therefore it can be started that conductance of weak electrolyte increases with dilution. The degree of dissociation of weak electrolyte needs to be calculated at particular concentration and hence dissociation constant can be calculated by applying Ostwald's law of dilution.

If degree dissociation of an electrolyte is considered to be 0.2, it means 20% of electrolyte is dissociated into ions and 80% of electrolyte will be in undissociated condition. Therefore, only 20% of electrolyte will give conductance value. The remaining 80% of electrolyte will not take part in conducting the electricity. As the conductivity value is proportional to number of dissociated molecules the degree of dissociation can be calculated by conductivity measurements.

Weak electrolytes like acetic acid, ammonium hydroxide, sodium acetate and ammonium chloride have low value of conductance at low dilution as they are partially ionized in the solution.

Observed conductivity:

The conductivity recorded by any conductivity cell is termed as observed conductivity.

Specific conductivity:

The specific conductivity of an electrolytic solution may be define as the conductance of a solution is placed between electrodes 1 cm apart and having cross-sectional area 1 sq.cm of each electrode. Specific conductivity means conductance of one cubic centimeter of electrolyte. The value of specific conductance depends upon the number of ions (concentration of ions) present in one cubic centimetre of solution.

The specific conductivity decreases with increasing dilution because number of ions per unite volume decreases.

The specific conductivity = cell constant x observed conductance

Equivalent conductivity (Λ_c):

It is the power of all ions produced by one gram equivalent of an electrolyte at any particular concentration.

$$\Lambda_c = K_c\ 1000/C$$

Where, Λ_c is equivalent conductivity, K_c is specific conductivity and C is the concentration.

The equivalent conductivity increases with increasing dilution and attains a limiting value at infinite dilution. The limiting value is called equivalent conductance at infinite dilution, denoted

by Λ_{∞}. The increase of equivalent conductivity is due to the increase of ions as a result of dissociation. As the ions get sufficient space to move, migration velocity increases.

Molecular conductivity:

Molecular conductance of a solution is the conductivity of all the ions produced by dissociation of one gram molecular weight of electrolyte when dissolved in Vml.

$$\mu_v = K_c \times 1000 / M$$

where, μ_v is the molecular conductivity.

Ionic conductance:

Weak electrolytes are partially ionized in solution and ionization increases with dilution. The conductivity is due to mobility of cations and anions. Hydrogen ions (H^+) is having high mobility than Na^+ and $CHCOO^-$ ions because it is lighter than sodium ion and acetate ion. It is not possible to calculate Λ_{∞} of weak electrolyte experimentally. Then can be calculated by applying Kolhraush law.

$$\Lambda_{\infty} = \text{Ionic conductivity of cation} + \text{Ionic conductivity of anion}$$

$$\Lambda_{\infty} = \lambda_c + \lambda_a$$

The molar ionic conductivity values are referred from literature.

Molar/Ionic conductivity values as 25 °C

Cation	λ^0_+ (S-cm^2/mol)	Anion	λ^0_- (S-cm^2/mol)
H^+	349.6	OH^-	199.1
Li^+	38.7	F^-	55.4
Na^+	50.10	Cl^-	76.35
K^+	73.50	Br^-	78.1
Rb^+	77.8	I^-	76.8
Cs^+	77.2	NO_2^-	71.8
Ag^+	61.9	NO_3^-	71.46
NH_4^+	73.5	ClO_3^-	64.6
Ethylammonium	47.2	ClO_4^-	67.3
Diethylammonium	42.0	IO_4^-	54.5
Triethylammonium	34.3	HCO_3^-	44.5
Tetraethylammonium	32.6	$H_2PO_4^-$	57
Tetra-*n*-butylammonium	19.5	HSO_3^-	50
Dimethylammonium	51.8	HSO_4^-	50
Trimethylammonium	47.2	$HC_2O_4^-$	40.2
Tetramethylammonium	44.9	$HCOO^-$	54.6
Piperidinium	37.2	CH_3COO^-	40.9
		$C_6H_5COO^-$	32.4

Exp. 1

Variation of conductance with concentration (strong &weak) electrolytes

Theory

The specific conductance· K of any conductor is defined as the (reciprocal of the specific resistance) that is

$$K=(L/aR) \quad (\Omega^{-1} cm^{-1})$$

Where R is the resistance, L is the length in centimeters and a is the cross sectional area in square centimeters of the conductor. The equivalent conductance by multiplying the specific conductance by the volume of the solution measured in cubic centimeters containing 1 g equivalent of electrolyte

$$\Lambda = K.v \quad (\Omega^{-1} cm^{2} eqt.^{-1})$$

Hence, if c is the concentration of the solution in gram equivalents per liter, v is equal to 1000/c , and

$$\Lambda = K\ 1000/c \quad (\Omega^{-1} cm^{2} eqt.^{-1})$$

Note that an equivalent is the associated with the transfer of one faraday of electricity. Hence, if one mole carries n faradays of electricity there will be n equitant in one mole, Only in the case where one mole is associated with the transfer of one faraday the mole and equivalent are identical.

Both specific and equivalent conductance of a solution vary with concentration .Specific conductance increases with increase in concentration .While equivalent conductance increases with dilution, a limiting value Λo being obtained at infinite dilution (zero concentration). The equivalent conductance at zero concentration of an electrolyte is not the same as the equivalent conductance of the pure solvent .it is the equivalent conductance of the solution when the ions of the electrolyte

are on infinite distance apart, and hence free from interaction.

For certain electrolytes, Kohlrausch suggested the empirical relationship

$$\Lambda = \Lambda o - b\sqrt{c}$$

Where b is a constant .These electrolytes are known as strong electrolytes and a plot of Λ vs. √c is linear for concentration up to 0.01 N .The equation is accurate to within a few percent ,however ,for concentrations of up to 0.1 N in the case of unit-univalent electrolytes .dilute aqueous solutions 0f many other electrolytes ,e.g. acetic acid ,do not give a linear graph .For

electrolytes , known as weak electrolytes, the value of the equivalent conductance increases rapidly as $\sqrt{c}$ approaches zero, and the graph cannot be accurately extrapolated to obtain a value of Λo .

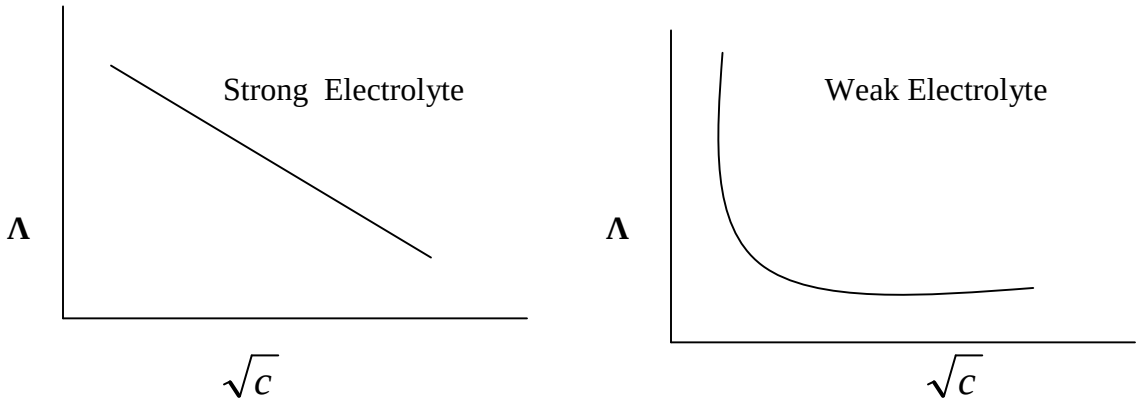
Procedure

1. Prepare 50 ml of 0.1 M (M/10) standard potassium chloride solution to be used for determining cell constant. The electrodes are washed with distilled water and the conductance of it 50ml recorded. And read conductance of KCl for eight diluted (0.01, 0.02, 0.03, 0.04, 0.05, 0.06, 0.07 and 0.08 molar solutions which prepared from 0.1M potassium chloride stock solution).
2. Prepare 50 ml of 0.1 M acetic acid. Read the conductance of 0.1 M acetic acid and for eight diluted (0.01, 0.02, 0.03, 0.04, 0.05, 0.06, 0.07 and 0.08 molar solutions which prepared from 0.1M acetic acid stock solution).

Calculation

1. Calculate the cell constant. κ = G **K** cm^{-1} , where **K** is the "cell constant" and G is the observed conductance, k is the specific conductivity for M/10 KCl at 25°C = 0.01288 mhos.
2. Equivalent conductance (Λ),

   Λ =149.82-93.85 $\sqrt{C}$ **+94.9c(1-0.2274** $\sqrt{C}$ **)**
3. Specific conductance (k), **K = Λ . C /1000 (Ω^{-1} cm^{-1})**
4. Molar conductance for strong and weak Electrolyte, Λm = k/C, where k is the specific constant.
5. Plote **Λ** vs. $\sqrt{C}$ thane from the obtained curve show the type of week electrolyte.



Exp.2

Hydrolysis constant by a conductimetric method

Theory

When the salt of a strong acid and weak base such as aniline chloride is dissolved in water ,hydrolysis occurs ,the cations of the salt interacting with the solvent to form the conjugate base of the salt and the acid of the solvent.

$$C_6H_5NH_3^+ Cl^- + H_2O \text{ ====== } C_6H_5NH_2 OH + H^+ Cl^-$$

$$(1-\sim) C \qquad \sim C \qquad \sim C$$

If fraction α of the salt hydrolysis , the hydrolysis constant is given by :-

$$K_h = \sim^2 C/ (1-\sim)$$

Where v is the volume in liters containing one mole of salt. the conductivity of the salt in which hydrolysis has occurred will be made up partly of the conductivity due to the ions produced by hydrolysis .

If excess of aniline is added in this solution the hydrolysis solution is suppresses and therefore conductivity mainly depends upon salt solution in presence of aniline and strong acid HCl formed during the hydrolysis.

For determination of degree of hydrolysis of the salt by conductivity method, following different measurements of conductivity are necessary.

i) The conductivity of aniline hydrochloride in water at certain concentration.

ii) The conductance of salt solution in presence of base i.e aniline.

iii) The conductivity of strong acid formed during the hydrolysis process.

Let ~ be the degree of hydrolysis. Therefore,

$$\Lambda v = (1-\sim) \Lambda' v + \sim \Lambda'' v$$

$$\Lambda v - \Lambda' v = \sim (\Lambda'' v - \Lambda' v)$$

$$\sim = \Lambda v - \Lambda' v / (\Lambda'' v - \Lambda' v)$$

Where Λv is the equivalent conductivity of aniline hydrochloride salt in water, $\Lambda' v$ is the equivalent conductivity of aniline hydrochloride in presence of aniline and $\Lambda'' v$ is the equivalent conductivity of HCL at that concentration.

Hydrolysis constant can be calculated by Ostwald's law of dilution.

$$\mathbf{K_h} = \sim^2 C / (1-\sim)$$

And it is possible to calculate the affinity constant of aniline K_b from the relationship .

$$\mathbf{K_h = (K_w / K_b)}$$

Procedure

1. Prepare 50 ml of 0.1 N (N/10) standard potassium chloride solution to be used for determining cell constant. The electrodes are washed with distilled water and the conductance of it 50ml recorded.
2. Prepare N/32 solution of aniline chloride in water take 50ml of this solution is placed in the conductivity cell and note down the conductivity at 25°C.
3. Prepare N/64 and N/128 solution from N/32 solution of aniline chloride by dilution, and read their conductivity.
4. Prepare also N/32, N/64 and N/128 solution of aniline chloride by using N/32 solution of aniline as a solvent and read their conductivity at 25°C.

Calculation

1. Calculate the cell constant. Specific constant = cell constant of KCl x observed conductance, the specific conductivity for M/10 KCl at 25°C = 0.01288 mhos.
2. Specific constant (Kc) for each solution.
3. Equivalent conductance for each solution by this relation, $\Lambda_v = K_c\ 1000/C$.
4. Degree of hydrolysis at each concentration, $\sim = \Lambda v - \Lambda' v / (\Lambda'' v - \Lambda' v)$.
5. Hydrolysis constant of aniline chloride, $K_h = \sim^2 C / (1-\sim)$.
6. Dissociation constant of aniline, $K_h = (K_w / K_b)$.

$\Lambda''v$ for hydrochloric acid at 25°C

N/32 393 Ω^{-1} cm^2 eqt.$^{-1}$

N/64 399 Ω^{-1} cm^2 eqt.$^{-1}$

N/128 401 Ω^{-1} cm^2 eqt.$^{-1}$

K_w for water at 25 °C=l.0l * 10^{-14} (g ions)^2 L^{-2}

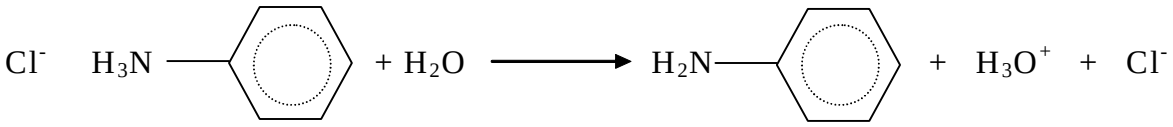
Exp.3

Hydrolysis constant by a pH method

Theory

The value of hydrogen concentrated predicate from hydrolysis of strong acid and weak base or from strong base and weak acid depend of the formation of pH or from acid value of solution:

$$\mathbf{pH = -\log[H^+]}$$



C(1-α) α C α C

Where α is degree of hydrolysis, and C concentration of solution.

$$\mathbf{K_h = (\sim^2 C) / (1-\sim)}$$

K_h is hydrolysis constant

Procedure

1. Prepare 50 ml of 0.1M of aniline hydrochloride and dilute from it 0.001 , 0.0001 and 0.00001 M .
2. Read pH of solutions after 20 min, when at hydrolysis is complete.

Calculation

1. calculation the hydrogen ion concentration at each stochiometric concentration by

$$\mathbf{pH = -\log[H^+]}$$

2. calculate degree of hydrolysis of each concentrated ion by

$$\mathbf{[H^+] = \sim C}$$

3. calculate the hydrolysis constant of each concentration by

$$\mathbf{K_h = (\sim^2 C) / (1-\sim)}$$

4. calculate ionic hydrolysis of aniline by

$$\mathbf{K_b = K_w / K_h}$$

Exp. 4

Dissociation constant of the weak acid (Conductometric method)

Theory

According to the so-called Ostwald dilution law, for a binary weak electrolyte, e.g. the weak acid HA.

$$HA + H_2O \text{ ======= } H_3O^+ + A^-$$

$$c(1 - \sim) \qquad \sim c \qquad \sim c$$

$$Kd = \frac{\sim 2.c}{(1 - \sim)}$$

Where a is the degree of dissociation, c is the concentration, and Kd is a constant sometimes referred to as the classical dissociation constant according to the Arrhenius dissociation theory, the degree of dissociation is given by the ratio of the equivalent conductance to the conductance at infinite dilution.

$$\alpha = \Lambda / \Lambda o$$

Hence, substituting for is the expression for the dissociation constant we have.

$$\mathrm{Kd} = \frac{\Lambda^2 \cdot \mathrm{c}}{\Lambda o(\Lambda o - \Lambda)}$$

$$\Lambda^2 \mathrm{c} = \mathrm{Kd}\, \Lambda o^2 - \mathrm{Kd}\, \Lambda o\, \Lambda$$

$$\Lambda . \mathrm{c} = \mathrm{Kd}\, \Lambda o^2 \cdot \frac{1}{\Lambda} - \mathrm{Kd}\, \Lambda o$$

$$\Lambda . c = Kd\left(\frac{\Lambda o^2 - \Lambda o}{\Lambda} \right)$$

A plot of $\Lambda . c$ vs. $1/\Lambda$ will thus give a straight line as shown in Fig.1 at

C= 0, OX = $1/\Lambda o$, that is $\Lambda o = 1/OX$. at $1/\Lambda = 0$, OY = $-Kd . \Lambda o$, therefore $(OX)(OY) = -Kd$

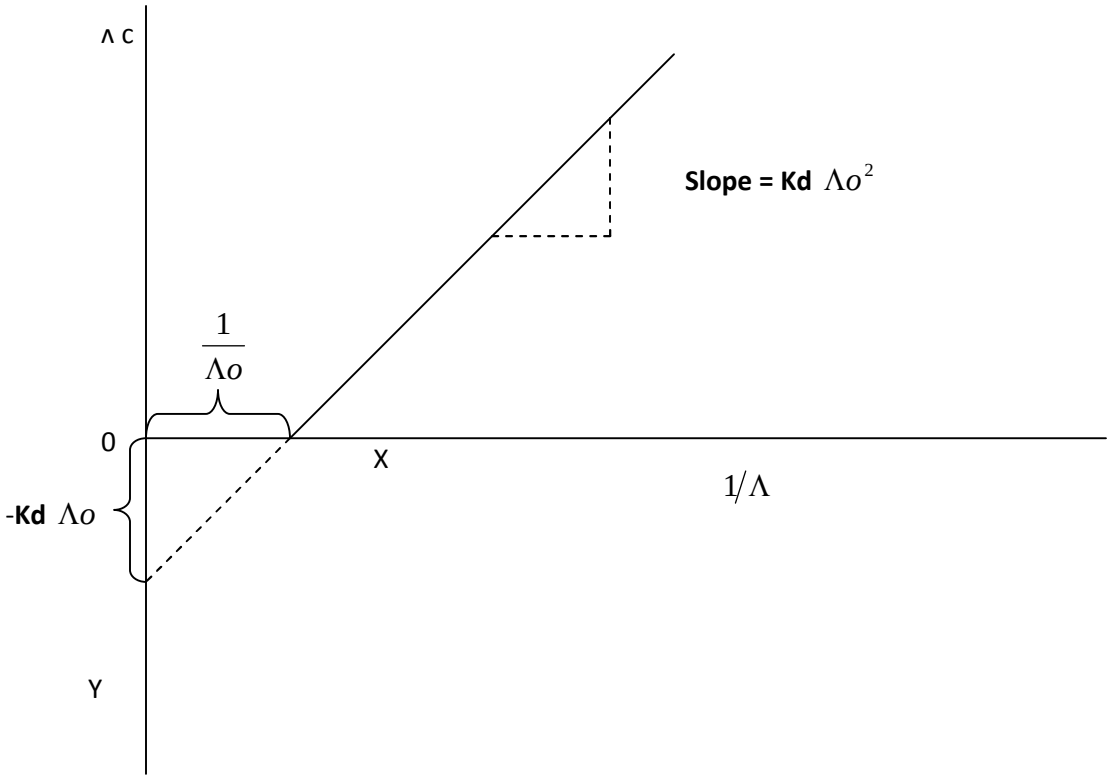
Where Kd is the classical dissociation constant

Procedure

1. Prepare 50 ml of 0.1 M (M/10) standard potassium chloride solution to be used for determining cell constant. The electrodes are washed with distilled water and the conductance of it 50ml recorded.
2. Prepare (0.01, 0.02, 0.03, 0.04, 0.05, 0.06, 0.08) solution from 0.1M of acetic acid stock solution. , and read their conductivity.

Calculation

1. Calculate the cell constant. Specific constant = cell constant of KCl x observed conductance, the specific conductivity for M/10 KCl at 25°C = 0.01288 mhos.
2. Specific constant (Kc) for each solution.
3. Equivalent conductance for each solution by this relation, $\mathbf{\Lambda_c = K_c\ 1000/C.}$
4. Plot of Λc vs.$1/\Lambda$.



Exp.5

Conductometric titration of strong, weak mixture acid with strong base

Theory

On the addition of a solution of a strong acid, there is initially a decrease in conductivity due to the replacement of the fast moving hydrogen ion by the slower alkali metal ion the conductivity subsequently increases again with the increase in concentration of hydroxyl ions when the titration is complete. Hence if the conductivity is plotted against the volume of base solution added the end point of the titration given by a Sharp inflection in the graph.

If strong base, however, is added to a solution of a weak acid, there is an initial gradual increase in the conductivity as the feebly dissociated weak electrolyte is converted into its salt, a stronger electrolyte. Subsequently at the end point a much more rapid conductivity increase occurs due to the presence of excess hydroxyl ions. If a mixture of strong acid and a weak acid is titrated by a strong base, an initial decrease in conductivity is followed by an increase, and then by a more rapid increase thereby giving two end points, see Figs.

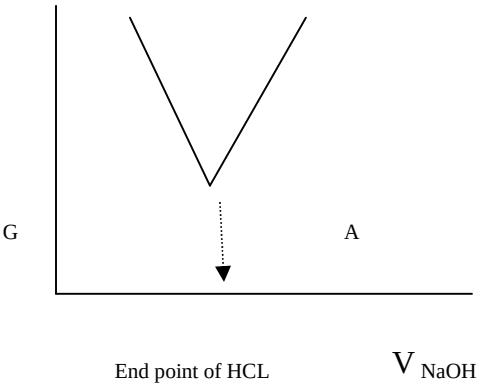
Procedure

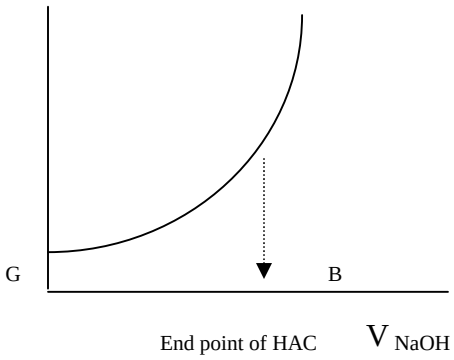
1. Prepare 0.1 M from each sodium hydroxide, hydrochloric acid and acetic acid in 50 ml.
2. Dilute 10 ml from hydrochloric acid in 50 ml and immersed in beaker to measured conductivity the 0.1 M sodium hydroxide solution is then run to acid solution and the conductance of the solution is measured after addition of 2,4,6,8,10 ……. ml of sodium hydroxide to after end point is present before the measurement of the conductance is made the solution should be thoroughly stirred.
3. Repeat this experiment using acetic acid instead of hydrochloric acid.
4. Prepare mixture of 5ml hydrochloric acid and 5ml acetic acid and diluted to 50 ml distilled water and repeat adding of sodium hydroxide and measure conductivity

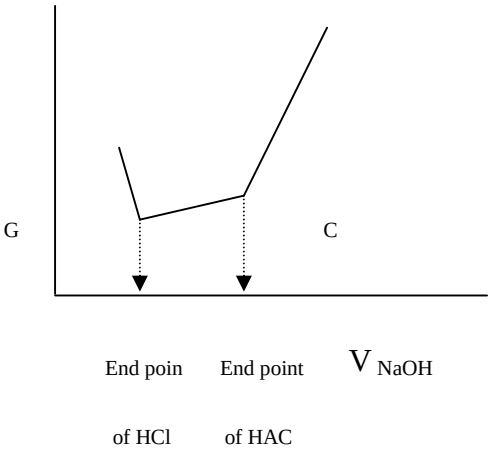
Calculation

1. Plotted graph conductivity vs. volume of sodium hydroxide for each acid and mixture.
2. Obtain from plotted the end point for each acid and mixture.
3. Determine the concentration of each acid using

$(N_1 . V_1)_{\text{of NaOH}} = (N_2 . V_2)_{\text{ of acid or mixture}}$







Exp.6

Solubility and solubility product from conductivity measurements

Theory

The Solubility product of an electrolyte may be determined from conductivity measurements. The conductance method is application for finding the concentration of the saturated solution provided that the salt is not hydrolysis and that solubility is not too high. It may then be assumed that the ions possess their limiting conductivities these must of course be known for the same temperature, measurement of the specific conductivity of the saturated solution leads to a value for the concentration.

Procedure

1. Prepare 50 ml of 0.1 M (M/10) standard potassium chloride solution to be used for determining cell constant. The electrodes are washed with distilled water and the conductance of it 50ml recorded.
2. Prepare silver chloride, silver bromide and silver iodide by preparation 0.1M from each potassium chloride, potassium bromide and potassium iodide in 50ml and prepare 0.1M of silver nitrate in 250 ml after this prepare mixed each 50ml of three sparingly soluble salts with 50ml from 0.1M silver nitrate, after this filtered each solution to measure conductivity.

Calculation

1. Calculate the cell constant. Specific constant = cell constant of KCl x observed conductance of 0.1 M KCl, the specific conductivity for M/10 KCl at 25°C = 0.01288 mhos.
2. Specific constant (Kc) for each solution.
3. Equivalent conductance for each solution by this relation, $\mathbf{\Lambda_c = K_c\ 1000/C.}$
4. Calculate equivalent conductance $\mathbf{\Lambda_\infty}$ for saturated solution by, $\mathbf{\Lambda_\infty = \lambda_c + \lambda_a}$.
5. Calculate solubility by, $\mathbf{\Lambda_\infty}$ = k /s x 1000.
6. Calculate the solubility product by, $Ksp = S^2$

Not / $\mathbf{\lambda_c}$ **and** $\mathbf{\lambda_a}$ it is equivalent ionic conductivities at finite dilution can be determined from references.

Exp. 7

Dissociation constant by the pH method

(pH titration curve)

Theory

One form of the glass electrode consists of a glass tube terminating in a thin walled glass bulb containing a platinium contact immersed in a buffer solution. Its potential (EG) when immersed in a buffer solution is dependent on the pH of the same test solution and is given by the equation

$$EG = \text{constant} + \frac{RT}{F} \ln aH+$$

Where aH+ is the activity of the hydrogen ions in the test solution. The precise potential of this electrode is not of great interest as it includes an asymmetry potential of this electrode is not off great interest as it includes an asymmetry potential (possibly due to strain at the glass solution interface)

This exists across the membrane even when the two solutions on either side of the membrane have identical activity values. In practice, however assembly shown below. Consequently this pH can be used as an indicator in acid – alkali titrations, there being a relatively large change in pH at the end point.

In this experiment the glass electrode is connected in series with saturated dip-type calomel electrode as shown:

Solution of constant pH | glass membrane| solution of unknown pH |Calomel electrode

The change in pH during the titration of sodium hydroxide with weak acid to determination the dissociation constant:

$$pH = pKa + \log \frac{b}{(a - b)}$$

When concentration of both base and acid, at equivalent point a = b and at halve equivalent point b = a/2, (pH)1/2 = pKa from (pH)1/2 calculate Ka value for acid.

Procedure

1. The electrode are washed in distilled water, dried with filter paper and washed again by successive immersions in the 0.1 M sodium hydroxide solution.
2. The electrodes are immersed in 25 ml of sodium hydroxide solution contained in a clean beaker.
3. This solution titrated with the 50 ml of 0.05 M acetic acid and the pH recorded after each addition of titrate, And stirring of the mixture before taking a pH reading.
4. On completion of the experiment the electrode is washed and immersed in distilled water.

Calculation

The flowing graphs are then plotted:-

a) pH vs. volume of acetic aced added Fig.1 and b)Δ (pH) / Δ (v) vs. volume of acetic acid added Fig.2

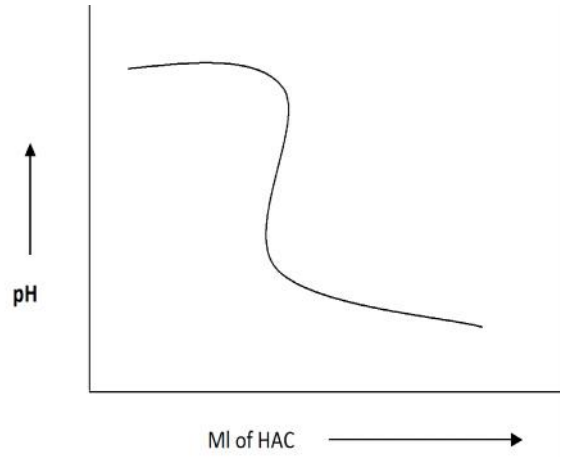


Fig.1

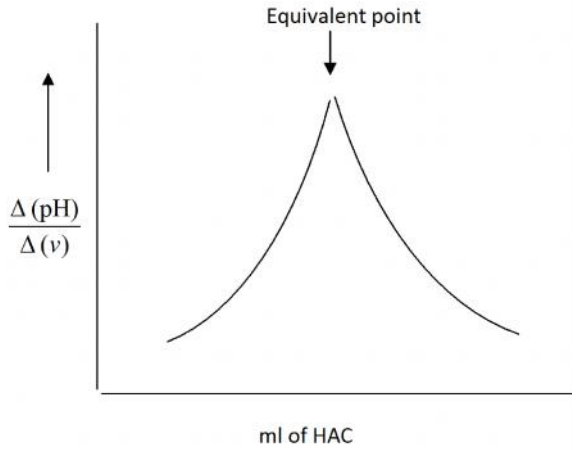


Fig.2

Safety Guidelines in the Chemistry Laboratory

The chemistry laboratory is a safe place to work if you are aware of important safety rules and if you are careful. You must be responsible for your own safety and for the safety of others. The safety rules given here will protect you and others from harm in the lab. While carrying out procedures in any of the **ChemLabs,** notice the safety symbols and caution statements. The safety symbols are explained in the chart on the next page.

1. Always obtain your teacher's permission to begin a lab.
2. Study the procedure. If you have questions, ask your teacher. Be sure you understand all safety symbols shown.
3. Use the safety equipment provided for you. Goggles and a safety apron should be worn when any lab calls for using chemicals.
4. When you are heating a test tube, always slant it so the mouth points away from you and others.
5. Never eat or drink in the lab. Never inhale chemicals. Do not taste any substance or draw any material into your mouth.
6. If you spill any chemical, wash it off immediately with water. Report the spill immediately to your teacher.
7. Know the location and proper use of the fire extinguisher, safety shower, fire blanket, first aid kit, and fire alarm.
8. Keep all materials away from open flames. Tie back long hair.
9. If a fire should break out in the classroom, or if your clothing should catch fire, smother it with the fire blanket or a coat, or get under a safety shower. **NEVER RUN.**
10. Report any accident or injury, no matter how small, to your teacher.

Follow these procedures as you clean up your work area.

1. Turn off the water and gas. Disconnect electrical devices.
2. Return materials to their places.
3. Dispose of chemicals and other materials as directed by your teacher. Place broken glass and solid substances in the proper containers. Never discard materials in the sink.
4. Clean your work area.
5. Wash your hands thoroughly after working in the laboratory.

First Aid in the Laboratory

Injury	Safe Response
Burns	Apply cold water. Call your teacher immediately.
Cuts and bruises	Stop any bleeding by applying direct pressure. Cover cuts with a clean dressing. Apply cold compresses to bruises. Call your teacher immediately.
Fainting	Leave the person lying down. Loosen any tight clothing and keep crowds away. Call your teacher immediately.
Foreign matter in eye	Flush with plenty of water. Use eyewash bottle or fountain.
Poisoning	Note the suspected poisoning agent and call your teacher immediately.
Any spills on skin	Flush with large amounts of water or use safety shower. Call your teacher immediately.

Safety Symbols

These safety symbols are used in laboratory and investigations in this book to indicate possible hazards. Learn the meaning of each symbol and refer to this page often. *Remember to wash your hands thoroughly after completing lab procedures.*

SAFETY SYMBOLS	HAZARD	EXAMPLES	PRECAUTION	REMEDY
DISPOSAL	Special disposal procedures need to be followed.	certain chemicals, living organisms	Do not dispose of these materials in the sink or trash can.	Dispose of wastes as directed by your teacher.
BIOLOGICAL	Organisms or other biological materials that might be harmful to humans	bacteria, fungi, blood, unpreserved tissues, plant materials	Avoid skin contact with these materials. Wear mask or gloves.	Notify your teacher if you suspect contact with material. Wash hands thoroughly.
EXTREME TEMPERATURE	Objects that can burn skin by being too cold or too hot	boiling liquids, hot plates, dry ice, liquid nitrogen	Use proper protection when handling.	Go to your teacher for first aid.
SHARP OBJECT	Use of tools or glassware that can easily puncture or slice skin	razor blades, pins, scalpels, pointed tools, dissecting probes, broken glass	Practice common-sense behavior and follow guidelines for use of the tool.	Go to your teacher for first aid.
FUME	Possible danger to respiratory tract from fumes	ammonia, acetone, nail polish remover, heated sulfur, moth balls	Make sure there is good ventilation. Never smell fumes directly. Wear a mask.	Leave foul area and notify your teacher immediately.
ELECTRICAL	Possible danger from electrical shock or burn	improper grounding, liquid spills, short circuits, exposed wires	Double-check setup with teacher. Check condition of wires and apparatus.	Do not attempt to fix electrical problems. Notify your teacher immediately.
IRRITANT	Substances that can irritate the skin or mucous membranes of the respiratory tract	pollen, moth balls, steel wool, fiberglass, potassium permanganate	Wear dust mask and gloves. Practice extra care when handling these materials.	Go to your teacher for first aid.
CHEMICAL	Chemicals that can react with and destroy tissue and other materials	bleaches such as hydrogen peroxide; acids such as sulfuric acid, hydrochloric acid; bases such as ammonia, sodium hydroxide	Wear goggles, gloves, and an apron.	Immediately flush the affected area with water and notify your teacher.
TOXIC	Substance may be poisonous if touched, inhaled, or swallowed.	mercury, many metal compounds, iodine, poinsettia plant parts	Follow your teacher's instructions.	Always wash hands thoroughly after use. Go to your teacher for first aid.
FLAMMABLE	Open flame may ignite flammable chemicals, loose clothing, or hair.	alcohol, kerosene, potassium permanganate, hair, clothing	Avoid open flames and heat when using flammable chemicals.	Notify your teacher immediately. Use fire safety equipment if applicable.
OPEN FLAME	Open flame in use, may cause fire.	hair, clothing, paper, synthetic materials	Tie back hair and loose clothing. Follow teacher's instructions on lighting and extinguishing flames.	Always wash hands thoroughly after use. Go to your teacher for first aid.

Eye Safety
Proper eye protection should be worn at all times by anyone performing or observing science activities.

Clothing Protection
This symbol appears when substances could stain or burn clothing.

Animal Safety
This symbol appears when safety of animals and students must be ensured.

Radioactivity
This symbol appears when radioactive materials are used.

Handwashing
After the lab, wash hands with soap and water before removing goggles

Physical Constants		
Quantity	**Symbol**	**Value**
Atomic mass unit	amu	1.6605×10^{-27}
Avogadro's number	N	6.022×10^{23} particles/mole
Ideal gas constant	R	8.31 L•kPa/mol•K 0.0821 L•atm/mol•K 62.4 mm Hg•L/mol•K 62.4 torr•L/mol•K
Mass of an electron	m_e	9.109×10^{-31} kg 5.485799×10^{-4} amu
Mass of a neutron	m_n	1.67492×10^{-27} kg 1.008665 amu
Mass of a proton	m_p	1.6726×10^{-27} kg 1.007276 amu
Molar volume of ideal gas at STP	V	22.414 L/mol
Normal boiling point of water	T_b	373.15 K 100.0°C
Normal freezing point of water	T_f	273.15 K 0.00°C
Planck's constant	h	$6.6260693 \times 10^{-34}$ J•s
Speed of light in a vacuum	c	2.997925×10^{8} m/s

Alphabetical Table of the Elements

Element	Symbol	Atomic number	Atomic mass
Actinium	Ac	89	227.027 8*
Aluminum	Al	13	26.981 539
Americium	Am	95	243.061 4*
Antimony	Sb	51	121.757
Argon	Ar	18	39.948
Arsenic	As	33	74.921 59
Astatine	At	85	209.987 1*
Barium	Ba	56	137.327
Berkelium	Bk	97	247.070 3*
Beryllium	Be	4	9.012 182
Bismuth	Bi	83	208.980 37
Bohrium	Bh	107	262*
Boron	B	5	10.811
Bromine	Br	35	79.904
Cadmium	Cd	48	112.411
Calcium	Ca	20	40.078
Californium	Cf	98	251.079 6*
Carbon	C	6	12.011
Cerium	Ce	58	140.115
Cesium	Cs	55	132.905 43
Chlorine	Cl	17	35.452 7
Chromium	Cr	24	51.996 1
Cobalt	Co	27	58.933 20
Copper	Cu	29	63.546
Curium	Cm	96	247.070 3*
Dubnium	Db	105	262*
Dysprosium	Dy	66	162.50
Einsteinium	Es	99	252.082 8*
Erbium	Er	68	167.26
Europium	Eu	63	151.965
Fermium	Fm	100	257.095 1*
Fluorine	F	9	18.998 403 2
Francium	Fr	87	223.019 7*
Gadolinium	Gd	64	157.25
Gallium	Ga	31	69.723
Germanium	Ge	32	72.61
Gold	Au	79	196.966 54
Hafnium	Hf	72	178.49
Hassium	Hs	108	265*
Helium	He	2	4.002 602
Holmium	Ho	67	164.930 32
Hydrogen	H	1	1.007 94
Indium	In	49	114.82
Iodine	I	53	126.904 47
Iridium	Ir	77	192.22
Iron	Fe	26	55.847
Krypton	Kr	36	83.80
Lanthanum	La	57	138.905 5
Lawrencium	Lr	103	260.105 4*
Lead	Pb	82	207.2
Lithium	Li	3	6.941
Lutetium	Lu	71	174.967
Magnesium	Mg	12	24.305 0
Manganese	Mn	25	54.938 05
Meitnerium	Mt	109	266*
Mendelevium	Md	101	258.098 6*
Mercury	Hg	80	200.59
Molybdenum	Mo	42	95.94
Neodymium	Nd	60	144.24
Neon	Ne	10	20.179 7
Neptunium	Np	93	237.048 2
Nickel	Ni	28	58.6934
Niobium	Nb	41	92 .906 38
Nitrogen	N	7	14. 006 74
Nobelium	No	102	259.100 9*
Osmium	Os	76	190.2
Oxygen	O	8	15.999 4
Palladium	Pd	46	106.42
Phosphorus	P	15	30.973 762
Platinum	Pt	78	195.08
Plutonium	Pu	94	244.064 2*
Polonium	Po	84	208.982 4*
Potassium	K	19	39.098 3
Praseodymium	Pr	59	140.907 65
Promethium	Pm	61	144.912 8*
Protactinium	Pa	91	231.035 88
Radium	Ra	88	226.025 4
Radon	Rn	86	222.017 6*
Rhenium	Re	75	186.207
Rhodium	Rh	45	102.905 50
Rubidium	Rb	37	85.467 8
Ruthenium	Ru	44	101.07
Rutherfordium	Rf	104	261*
Samarium	Sm	62	150.36
Scandium	Sc	21	44.955 910
Seaborgium	Sg	106	263*
Selenium	Se	34	78.96
Silicon	Si	14	28.085 5
Silver	Ag	47	107.868 2
Sodium	Na	11	22.989 768
Strontium	Sr	38	87.62
Sulfur	S	16	32.066
Tantalum	Ta	73	180.947 9
Technetium	Tc	43	97.907 2*
Tellurium	Te	52	127.60
Terbium	Tb	65	158.925 34
Thallium	Tl	81	204.383 3
Thorium	Th	90	232.038 1
Thulium	Tm	69	168.934 21
Tin	Sn	50	118.710
Titanium	Ti	22	47.88
Tungsten	W	74	183.85
Uranium	U	92	238.028 9
Vanadium	V	23	50.941 5
Xenon	Xe	54	131.29
Ytterbium	Yb	70	173.04
Yttrium	Y	39	88.905 85
Zinc	Zn	30	65.39
Zirconium	Zr	40	91.224

* The mass of the isotope with the longest known half-life.